



Experiment 54

Molar Volume and the Universal Gas Constant

Problem

How can the value of the universal gas constant be verified experimentally?

Introduction

There are a number of gas laws described in Chapter 13. Some, like the laws of Avogadro, Boyle and Charles, concern how one of the four variables – pressure, volume, absolute temperature, and number of moles – varies when another is changed, with two of the four being held constant. Two other relationships are of a more general nature. One, the Combined Gas Law, represents a combining of the simpler laws. Given that volume is inversely proportional to pressure (Boyle's Law), and directly proportional to temperature (Charles' Law), you can write an expression that incorporates both sets of observations:

$$P_1V_1/T_1 = P_2V_2/T_2$$

Going one step further and introducing the law of Avogadro, which says that the volume occupied by a gas is directly proportional to the number of moles of gas in the system, gives the expression

$$P_1V_1/n_1T_1 = P_2V_2/n_2T_2$$

In which P_1 , V_1 , n_1 and T_1 refer to an initial set of conditions and P_2 , V_2 , n_2 and T_2 refer to the final conditions, after the change has occurred. This "combined" gas law allows you to do calculations involving changes in any or all of the four variables.

What the equation tells us is that the value of the expression, PV/nT , does not change; if that is the case, then we can determine a value for PV/nT that should be valid under any set of conditions as long as the sample remains a gas. In other words we can state that

$$PV/nT = \text{constant} = R$$

This relationship ($PV/nT = R$) can be rearranged to the form known as the Ideal Gas Law:

$$PV = nRT$$

The constant, R , is known as the *universal gas constant*. One of your objectives in this experiment is to determine the value of R experimentally. Notice that the units of R must reflect pressure times volume, divided by number of moles times temperature. R is most often expressed in units of L atm/mol K ("liter-atmospheres per mole-kelvin"). In this experiment you will calculate the value of R for each of three trials, as well as an average result which will then be used to determine your percentage error.

You will first determine the molar volume of a gas. You will then use the molar volume at laboratory conditions to determine what volume one mole of gas would occupy at STP (0°C and one atmosphere pressure). You can make this conversion by using the combined gas law. Once that is done, you will use your experimental values of P , V , n , and T to calculate an experimental value for R . As with the molar volume calculation, you will determine individual values for each trial, along with an average value, which you will compare with the accepted value of R : $0.0821 \text{ L atm/mol}\cdot\text{K}$.

Prelaboratory Assignment

- ✓ Read the **Introduction and Procedure** before you begin.
 - ✓ Answer the Prelaboratory Questions.
1. Explain how the mass of a piece of magnesium ribbon may be calculated from its length and the mass of a 100.0-cm strip of the same type of ribbon.
 2. What two physical properties of hydrogen gas make it possible for you to collect it by displacement of water in your graduated cylinder?
 3. Use Dalton's Law of Partial Pressures to explain why the pressure of hydrogen gas in the cylinder will be less than the observed barometric pressure in the laboratory. How will you determine the pressure of the hydrogen you produce?
 4. Calculate the mass of baking soda, NaHCO_3 , needed to neutralize 3.0 mL of 3.0 M hydrochloric acid, $\text{HCl}(aq)$. Show your work; the answer alone is not enough. (**Hint:** the amount needed is less than 1 gram.)
 5. The accepted value for the universal gas constant is $0.0821 \text{ L atm/mol}\cdot\text{K}$. What would it be if the pressure was measured in torr and the volume in milliliters? Show your calculations.

Materials

Apparatus

10-mL graduated cylinder
#00 1-hole rubber stopper
Copper wire, 10 cm length
Thermometer
Safety goggles
Lab apron

Reagents

Magnesium ribbon, 3 pcs, $\sim 0.9 \text{ cm}$
3 $M \text{ HCl}(aq)$
 $\text{NaHCO}_3(s)$ (baking soda)

Safety



1. Wear safety goggles and a lab apron at all times in the laboratory.
2. Hydrochloric acid is corrosive to skin and clothing. Clean up all spills thoroughly.
Note: If acid spills on the lab bench, use a bit of baking soda to neutralize it before cleaning up.
Do not neutralize acid that spills on skin or clothing; flood the affected area with water.
3. If you are using a mercury thermometer; be very careful. Mercury vapor is very poisonous. If the thermometer breaks, notify your teacher right away.

Procedure

1. Fill a 400-mL beaker about three-fourths full of tap water; the water should be at or near room temperature.
2. Obtain a short (0.80-1.00 cm) piece of magnesium ribbon, then measure and record its length to the nearest 0.01 cm. Use a 10-cm piece of copper wire to make a cage for your magnesium, by folding the magnesium over the wire then rolling the wire around the magnesium. Fit the wire cage into a 1-hole #00 rubber stopper. The cage should be about 2-3 cm from the small end of the stopper to hold the magnesium in place. See Figure 1.

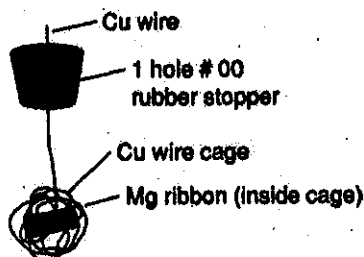


Figure 1

3. Carefully pour about 3 mL of 3 M HCl(aq) into a 10-mL graduated cylinder. Using a wash bottle or beaker, carefully add distilled or deionized water to the graduated cylinder until it is completely full. Try to direct the water down the side of the graduated cylinder to prevent mixing of the water and the acid. Insert the stopper assembly (see Fig. 1, above) into the top of the graduated cylinder. Water should escape through the hole in the stopper; if it does not, remove the stopper and carefully add more water, then replace the stopper assembly. This will keep air from being trapped in the graduated cylinder.
4. Place your finger over the hole in the stopper and invert the graduated cylinder, lowering it into the beaker of water. Remove your finger when the stopper is below the level of water in the beaker. The hydrochloric acid is more dense than pure water, so it will slowly sink toward the stopper and the magnesium; observe and record evidence of reaction. (See Figure 2.)

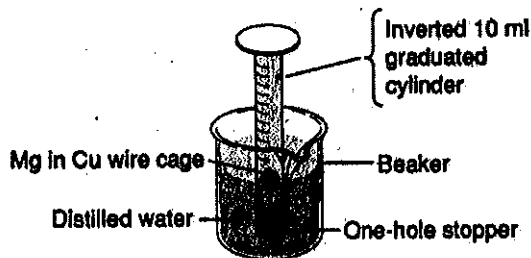


Figure 2

5. When the reaction is complete, allow the system to stand for two or three minutes, tapping the sides of the graduated cylinder to dislodge any gas bubbles that may be clinging to the glass wall. Make sure that there are no little pieces of unreacted magnesium on the wall of the graduated cylinder. If a small piece remains, gently shake the graduated cylinder up and down to wash the metal back into the acid solution, allowing it to finish reacting. Be careful not to lift the graduated cylinder completely out of the water in the beaker.

6. Adjust the position of the graduated cylinder so that the water levels inside and out are even; this ensures that the total pressure on the gases inside the graduated cylinder is the same as the barometric pressure in the room. Record the volume of gas trapped in the graduated cylinder, and the temperature of the water near the mouth of the graduated cylinder. The temperature of the escaping solution may be assumed to be the same as the temperature of the trapped gases. Enter these data in the Data Table.

Cleaning Up



1. Take apart the apparatus. Dispose of the copper wire in the solid waste container or as your teacher directs.
2. Use baking soda to neutralize the acidic solution remaining in the beaker. The neutralized solution can be flushed down the drain safely. As part of the Prelaboratory Assignment, you calculated the mass of baking soda needed. Use a plastic spoon to add approximately that amount of the solid (a little at a time to minimize foaming) to the beaker. Stir well and flush the solution down the drain.
3. Clean all glassware and return it to its proper location.
4. Wash your hands before leaving the laboratory.

Analysis and Conclusions

Complete the **Analysis and Conclusions** section for this experiment either on your Report Sheet or in your lab report as directed by your teacher.

All of the calculations are to be shown for your first trial; you may simply report the results for the other two. Enter the results for all trials in a Summary Table.

1. Calculate the mass of magnesium for each trial, using the mass of 1.00 m of magnesium ribbon supplied by your instructor.
2. Calculate the number of moles of magnesium used. This is the same as the number of moles of hydrogen generated. (Why?)
3. Because you collected hydrogen over water, a small portion of the gas in the cylinder at the end of the reaction is water vapor; we say the hydrogen gas is "wet." The amount of water that evaporates is dependent only on the temperature, so it is a simple matter to determine the *partial pressure* of the water vapor in the graduate. Use the table of vapor pressures, found in Appendix A of your lab manual, to find the pressure due to water in the graduated cylinder. Subtract this value from the barometric pressure to get the pressure exerted by the "dry" hydrogen (hydrogen without the water vapor). Be sure to report the pressure to the correct degree of precision. Enter the results in Summary Table 1.

4. You have measured the volume occupied by a very small fraction of a mole of hydrogen, under a specific set of conditions of pressure and temperature. The volume occupied by one mole of gas is called the **molar volume** of the gas, and it is the same for all gases (behaving ideally) at a particular pressure and temperature. For each trial:
 - a. Calculate the volume that 1.00 mole of hydrogen would occupy at your experimental temperature and pressure (called "laboratory conditions"). Record your answers in the Summary Table.
 - b. Use the combined gas law to calculate the volume that 1.00 mole of hydrogen would occupy at 1.00 atm and 273 K (Standard Temperature and Pressure, STP).
5. Determine the value of PV/nT for each trial. P is the pressure of dry hydrogen, V is the volume of gas collected, T is the Kelvin temperature, and n is the number of moles of hydrogen generated.
6. Determine the average of your three values for PV/nT . Also determine the deviation for each of your three trials and the average deviation.
7. Your average for PV/nT represents your experimental value for the universal gas constant, R . Calculate the percent error in your determination of the value of R .